

PROSPERITY ACADEMY

AS CHEMISTRY 9701

Crash Course

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**ATOMS, MOLECULES
AND STOICHIOMETRY**

COMPLETE NOTES



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Atoms, molecules and Stoichiometry

1. Relative atomic mass (A_r):- The weighted average mass of all naturally occurring isotopes of an element on a scale where 1 atom of Carbon-12 has a mass of 12 amu
2. Isotopes:- Atoms of the same element having same number of protons but different number of neutrons.
3. Relative isotopic mass:- Mass of a particular isotope of an element on a scale where 1 atom of Carbon-12 has a mass of 12 amu
4. Relative Molecular mass (M_r):- The mass of one molecule of an element or compound on a scale where 1 atom of Carbon-12 has a mass of 12 amu
5. Relative formula mass:- The mass of one formula unit of a compound on a scale where 1 atom of Carbon-12 has a mass of 12 amu
6. Mole:- Amount of substance containing same number of particles as 12g of Carbon-12 / or 6.02×10^{23} particles
7. Molar mass:- The mass of 1 mole of a substance in g mol^{-1} .
8. Solution Concentration:- The amount of solute (in mol) dissolved in 1 dm^3 of solution.

Formulas and Calculations:-

$$\text{mol} = \frac{\text{Mass}}{M_r / A_r}$$

$$\text{No. of particles (N)} = \text{mol} \times 6.02 \times 10^{23}$$

Q. Calculate the number of particles in:-

a) 35.5 g of Cl_2

$$\text{mol} = \frac{35.5}{71} = 0.5 \text{ mol}$$

$$N = 0.5 \times 6.02 \times 10^{23} = 3.01 \times 10^{23} \text{ particles}$$

b) 336 g of iron

$$\text{mol} = \frac{336}{55.8} = 6.022 \text{ mol}$$

$$N = 6.022 \times 6.02 \times 10^{23} = 3.62 \times 10^{23}$$

Q. What mass of Aluminium contains

a) 2×10^{23} atoms

$$\text{mol} = \frac{2 \times 10^{23}}{6.02 \times 10^{23}} = 0.3322 \text{ mol}$$

$$\text{mass} = 0.3322 \times 27 = 8.97 \text{ g} \approx 9 \text{ g}$$

b) 6×10^{20} atoms

$$\text{mol} = \frac{6 \times 10^{20}}{6.02 \times 10^{23}} = 9.967 \times 10^{-4}$$

$$\text{mass} = 9.967 \times 10^{-4} \times 27 = 0.0269 \approx 0.03 \text{ g}$$

- One mole of gas occupies $24 \text{ dm}^3 / 24000 \text{ cm}^3$ of volume at r.t.p.
- $\text{Volume at r.t.p. (V)} = \text{mol} \times 24 \text{ dm}^3 / 24000 \text{ cm}^3$

Q. Calculate the volume occupied by 0.40 mol of NO_2 at r.t.p.

$$0.40 \times 24 \text{ dm}^3 = 9.6 \text{ dm}^3$$

Q. Calculate the mass of CO_2 present in 500 cm^3 at r.t.p.

$$\text{mol} = \frac{500}{24000} = \frac{1}{48} \text{ mol} \quad \Rightarrow \quad \text{Mass} = \frac{1}{48} \times (12 + 32) = 0.91667 \text{ g} \approx 0.9 \text{ g}$$

Q. Calculate the volume occupied by 32g of CO_2 at r.t.p.

$$\text{mol} = \frac{32}{(12 + 32)} = 0.72727 \quad V = 0.72727 \times 24 = 17 \text{ dm}^3$$

- $\text{concentration (mol dm}^{-3}\text{)} = \text{moles} / \text{Volume (dm}^3\text{)}$

Q. What is the concentration of CH_3COOH solution, containing 12g of CH_3COOH in 250 cm^3 of solution?

$$\text{mol} = \frac{12}{60} = 0.2 \text{ mol}$$

$$\text{concentration} = 0.2 / (250/1000) = 0.80 \text{ mol dm}^{-3}$$

Q. Calculate the mass of anhydrous copper (II) sulfate in 55 cm^3 of a 0.20 mol dm^{-3} solution of copper (II) sulfate.

$$\text{mol} = \text{conc} \times \text{volume} = 0.20 \times \frac{55}{1000} = 0.011 \text{ mol}$$

$$\text{mass} = \text{mol} \times \text{Mr / Ar (CuSO}_4\text{)}$$

$$0.011 \times [63.5 + 32.1 + 4(16)] = 1.7556 \text{ g} \approx 1.8 \text{ g}$$

Empirical and Molecular formulae
simplest ratio formula \rightarrow exact formula

Q. 9.8g of an oily liquid was found to contain 0.2g of hydrogen, 3.2g of sulfur and 6.4g of oxygen
Find the empirical formula of the liquid.

	H	S	O
mass	0.2	3.2	6.4
moles	$\frac{0.2}{1}$	$\frac{3.2}{32.1}$	$\frac{6.4}{16}$
	$= \frac{0.2}{0.09968}$	$= \frac{0.09968}{0.09968}$	$= \frac{0.4}{0.09968}$

simplest ratio = 2 : 1 : 4

Empirical formula :- H_2SO_4

- Molecular formula = (Empirical formula)_n

- $n = \text{Molar mass of molecule} / \text{Molar mass of empirical formula}$

Q. An organic compound contains 58.8% carbon, 9.8% hydrogen and 31.4% oxygen. Its molar mass is 102 g mol⁻¹. Find the empirical and molecular formulae of the compound.

	C	:	H	:	O		$n = \frac{102}{102} = 1$
mass	58.8	:	9.8	:	31.4		
moles	$\frac{58.8}{12}$ = 4.9	:	$\frac{9.8}{1}$ = 9.8	:	$\frac{31.4}{16}$ = 1.9625		Molecular formula = C ₅ H ₁₀ O ₂
simplest ratio	$\frac{4.9}{1.9625}$:	$\frac{9.8}{1.9625}$:	$\frac{1.9625}{1.9625}$		
	2.5 × 2	:	5 × 2	:	1 × 2		
	5	:	10	:	2		

Empirical formula:- C₅H₁₀O₂ (mass:- 102)

Stoichiometry:- Titration Examples

4 Compound R is a weak diprotic (dibasic) acid which is very soluble in water.

(a) A solution of R was prepared which contained 1.25g of R in 250cm³ of solution. When 25.0cm³ of this solution was titrated with 0.100 mol dm⁻³ NaOH, 21.6cm³ of the alkali were needed for complete reaction.

(i) Using the formula H₂X to represent R, construct a balanced equation for the reaction between H₂X and NaOH.



(ii) Use the data above to calculate the amount, in moles, of OH⁻ ions used in the titration.

$$\text{mol} = \text{conc} \times \text{vol} = 0.100 \times \frac{21.6}{1000} = 2.16 \times 10^{-3} \text{ mol}$$

(iii) Use your answers to (i) and (ii) to calculate the amount, in moles, of R present in 25.0 cm³ of solution.

$$\frac{2.16 \times 10^{-3}}{2} = 1.08 \times 10^{-3} \text{ mol}$$

(iv) Calculate the amount, in moles, of R present in 250 cm³ of solution.

$$\begin{array}{l} 25: 1.08 \times 10^{-3} \\ 250: x \end{array} \Rightarrow x = \frac{1.08 \times 10^{-3} \times 250}{25} = 0.0108$$

(v) Calculate M_r of R.

$$M_r = \frac{\text{mass}}{\text{mol}} = \frac{1.25}{0.0108} = 115.7 \approx \boxed{116g}$$

6 25.0 cm³ of sulfuric acid solution reacts with 36.2 cm³ of 0.225 mol dm⁻³ sodium hydroxide solution. The concentration of the acid is

A $\frac{36.2 \times 0.225}{25.0}$ B $\frac{2 \times 36.2 \times 0.225}{25.0}$ C $\frac{36.2 \times 0.225}{2 \times 25.0}$ D $\frac{25.0}{2 \times 36.2 \times 0.225}$



$$\begin{aligned} \text{mol of NaOH} &= \text{conc} \times \text{vol} = 0.225 \times \frac{36.2}{1000} \\ &= 8.145 \times 10^{-3} \end{aligned}$$

$$\text{mol of } H_2SO_4 = \frac{1}{2} \times 8.145 \times 10^{-3} = 4.0725 \times 10^{-3}$$

$$\text{conc} = \frac{4.0725 \times 10^{-3}}{25/1000} = 0.1629$$

Stoichiometry:- Gas volumes

For only gas reactions, the mole ratio is equivalent to volume ratio

50 cm³ of propane was mixed in 600 cm³ of oxygen (an excess) and ignited.

What is the volume of carbon dioxide produced?

How much oxygen remains unused?



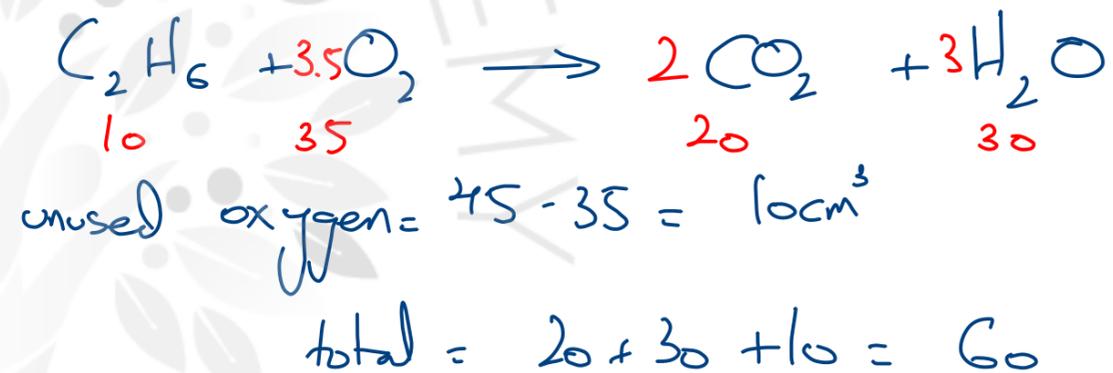
1) 150 cm³ of CO₂ is produced

2) 600 - 250 = 350 cm³ of O₂ remains

4 10 cm³ of ethane is burned in 45 cm³ of oxygen at a pressure of 101 kPa and a temperature of 200 °C. Complete combustion takes place.

What is the total volume of gas present when the reaction is complete, measured under the same conditions?

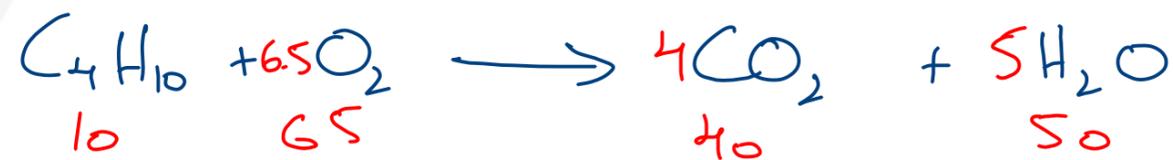
- A 30 cm³ B 50 cm³ C 55 cm³ **D 60 cm³**



6 In this question you should assume air contains 21% oxygen.

What is the minimum volume of air required to ensure complete combustion of 10 cm³ of butane gas, under room conditions?

- A 14 cm³ B 27 cm³ C 65 cm³ **D 310 cm³**

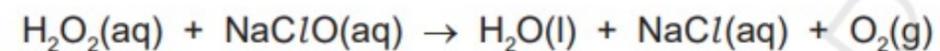


$$x \times \frac{21}{100} = 65 = 309.5 \approx 310\text{cm}^3$$

Stoichiometry:- Tougher Questions

(e) The concentration of NaClO in bleach **S** is $x \text{ g dm}^{-3}$.

NaClO reacts with $\text{H}_2\text{O}_2(\text{aq})$ as shown.



A 5.00 cm^3 sample of **S** completely reacts with $\text{H}_2\text{O}_2(\text{aq})$. The volume of $\text{O}_2(\text{g})$ produced is 24.0 cm^3 under room conditions.

Assume that only the NaClO in **S** reacts with $\text{H}_2\text{O}_2(\text{aq})$.

Calculate x . Show your working.

$$n \text{ of } \text{O}_2 = \frac{24}{24000} = 1 \times 10^{-3} \text{ mol}$$

$$n \text{ of NaClO} = 1 \times 10^{-3} \text{ mol}$$

$$\text{mass} = \text{mol} \times M_r/A_r$$

$$1 \times 10^{-3} \times (23 + 35.5 + 16) \\ = 0.0745 \text{ g}$$

$$\text{concentration} = \frac{0.0745}{5 \times 10^{-3}} \\ = 14.9 \text{ g dm}^{-3}$$

$$x = \dots\dots\dots 14.9 \dots\dots\dots \text{ g dm}^{-3} \\ [3]$$

- 2 Chile saltpetre is a mineral found in Chile and Peru, and which mainly consists of sodium nitrate, NaNO_3 . The mineral is purified to concentrate the NaNO_3 which is used as a fertiliser and in some fireworks.

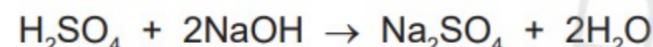
In order to find the purity of a sample of sodium nitrate, the compound is heated in $\text{NaOH}(\text{aq})$ with Devarda's alloy which contains aluminium. This reduces the sodium nitrate to ammonia which is boiled off and then dissolved in acid.



The ammonia gas produced is dissolved in an excess of H_2SO_4 of known concentration.



The amount of unreacted H_2SO_4 is then determined by back-titration with NaOH of known concentration.



- (a) A 1.64 g sample of impure NaNO_3 was reacted with an excess of Devarda's alloy.

The NH_3 produced was dissolved in 25.0 cm^3 of $1.00 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$.

When all of the NH_3 had dissolved, the resulting solution was titrated with $\text{NaOH}(\text{aq})$.

For neutralisation, 16.2 cm^3 of $2.00 \text{ mol dm}^{-3} \text{ NaOH}$ were required.

- (i) Calculate the amount, in moles, of H_2SO_4 present in the 25.0 cm^3 of $1.00 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$.

$$n = c \times V = 1 \times 25 \times 10^{-3} = 0.025$$

- (ii) Calculate the amount, in moles, of NaOH present in 16.2 cm^3 of $2.00 \text{ mol dm}^{-3} \text{ NaOH}$.

$$n = c \times V = 2 \times 16.2 \times 10^{-3} = 0.0324$$

- (iii) Use your answer to (ii) to calculate the amount, in moles, of H_2SO_4 that reacted with 16.2 cm^3 of $2.00 \text{ mol dm}^{-3} \text{ NaOH}$.

$$= \frac{0.0324}{2} = 0.0162$$

- (iv) Use your answers to (i) and (iii) to calculate the amount, in moles, of H_2SO_4 that reacted with the NH_3 .

$$0.025 - 0.0162 = 8.80 \times 10^{-3}$$

- (v) Use your answer to (iv) to calculate the amount, in moles, of NH_3 that reacted with the H_2SO_4 .

$$2 \times 8.80 \times 10^{-3} = 0.0176 \text{ mol}$$

- (vi) Use your answer to (v) to calculate the amount, in moles, of NaNO_3 that reacted with the Devarda's alloy.

$$= 0.0176 \text{ mol}$$

- (vii) Hence calculate the mass of NaNO_3 that reacted.

$$\text{mass} = \text{mol} \times M_r / A_r = 0.0176 \times (23 + 14 + 3(16)) = 1.496 \text{ g}$$

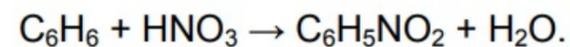
- (viii) Use your answer to (vii) to calculate the percentage by mass of NaNO_3 present in the impure sample.

Write your answer to a suitable number of significant figures.

$$\frac{1.496}{1.64} \times 100 = 91.2 \%$$

Percentage yield :-

Benzene, C_6H_6 , reacts with concentrated nitric acid at $40^\circ C$ to make nitrobenzene, $C_6H_5NO_2$, as follows:



A student used 10.0 g benzene and obtained 9.0 g nitrobenzene.

What percentage yield was produced?

$$n \text{ of } C_6H_6 = \frac{10}{78} = \frac{5}{39} \text{ moles}$$

$$n \text{ of } C_6H_5NO_2 = \frac{5}{39} \text{ moles}$$

$$\text{mass of } C_6H_5NO_2 = \frac{5}{39} \times 123 = 15.77 \text{ g}$$

$$\frac{9}{15.77} \times 100 = 57\%$$

Water of Crystallization:-

(b) In the reaction described in (a)(i), a student uses 17.43g of $\text{CuSO}_4 \cdot y\text{H}_2\text{O}$. By further titration of the reaction products the student concludes that the total amount of CuSO_4 in the sample is 0.0982 mol.

Use the *Data Booklet* to complete the table to calculate the value of y , where y is an integer. Show your working.

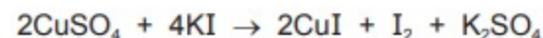
mass of 0.0982 mol CuSO_4	$0.0982 \times [63.5 + 32.1 + 4(16)]$ $= 15.672$ 15.67 g
amount of H_2O in 17.43g of $\text{CuSO}_4 \cdot y\text{H}_2\text{O}$	$17.43 - 15.67 = 1.757$ g $\text{mol} = \frac{1.757}{18} =$ 0.0978 mol H_2O
value of y	$\frac{0.0978}{0.0982} = 1$ $y =$ 1

[4]

[Total: 9]

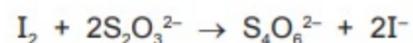
(d) A student is provided with a sample of hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, and is asked to determine the value of x .

The student dissolves a sample of the hydrated copper(II) sulfate in water and adds it to an excess of aqueous potassium iodide to make a total volume of 250.0 cm^3 of solution.



The amount of iodine produced during this reaction is found by titrating a sample of this solution with sodium thiosulfate solution.

25.0 cm^3 of the iodine-containing solution requires 20.0 cm^3 of 0.10 mol dm^{-3} sodium thiosulfate solution.



(i) Calculate the amount, in mol, of copper(II) sulfate present in the original sample of hydrated copper(II) sulfate.

Show your working.

$$n \text{ of } \text{S}_2\text{O}_3^{2-} = c \times V = 0.10 \times 20 \times 10^{-3} = 2 \times 10^{-3}$$

$$n \text{ of } \text{I}_2 = 1 \times 10^{-3}$$

$$\text{I}_2 \text{ in } 250 \text{ cm}^3 = \frac{1 \times 10^{-3} \times 250}{25} = 0.01$$

$$n \text{ of } \text{CuSO}_4 = 2 \times 0.01 = 0.02$$

amount of copper(II) sulfate = 0.02 mol [2]

(ii) A total of 7.98 g of CuSO_4 is present in 10.68 g of $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$.

Complete each row of the table to calculate the value of x , where x is an integer.

[M_r : CuSO_4 , 159.6]

amount of CuSO_4 in 10.68 g of $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$	$n = \frac{7.98}{63.5 + 32.1 + 4(16)} = 0.05$ <u>0.05</u> mol
amount of H_2O in 10.68 g of $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$	$10.68 - 7.98 = 2.7 \text{ g}$ $n = \frac{2.7}{18} = 0.15$ <u>0.15</u> mol
value of x	$x = \frac{0.15}{0.05}$ <u>3</u>

[3]

(d) Mohr's salt, $(\text{NH}_4)_2\text{Fe}(\text{SO}_4)_2 \cdot x\text{H}_2\text{O}$, is the hydrated form of ammonium iron(II) sulfate.

x represents the number of moles of water in 1 mole of the salt.

A student wanted to determine the value of x . 0.784 g of the hydrated salt was dissolved in water and this solution was acidified.

All of the solution was titrated with $0.0200 \text{ mol dm}^{-3}$ potassium manganate(VII). 20.0 cm^3 of this potassium manganate(VII) solution was required for complete reaction with the Fe^{2+} ions.

(i) Use changes in oxidation numbers to balance the equation for the reaction taking place.



(ii) State the role of the Fe^{2+} ions in this reaction.

Explain your answer.

(iii) Calculate the amount, in moles, of manganate(VII) ions that reacted.

$$n = c \times V = 0.02 \times 20 \times 10^{-3} = 4 \times 10^{-4}$$

amount = 4.00×10^{-4} mol [1]

(iv) Calculate the amount, in moles, of Fe^{2+} ions in the sample of the salt.

$$4.00 \times 10^{-4} \times 5 = 2.00 \times 10^{-3}$$

amount = 2.00×10^{-3} mol [1]

(v) Calculate the relative formula mass of $(\text{NH}_4)_2\text{Fe}(\text{SO}_4)_2 \cdot x\text{H}_2\text{O}$.

$$M_r / A_r = \frac{\text{mass}}{\text{mol}} = \frac{0.784}{2.00 \times 10^{-3}} = 392$$

relative formula mass = 392 [1]

(vi) Calculate the value of x .

$$392 = 28 + 8 + 55.8 + 64.2 + 8(16) + 2x + 16x$$

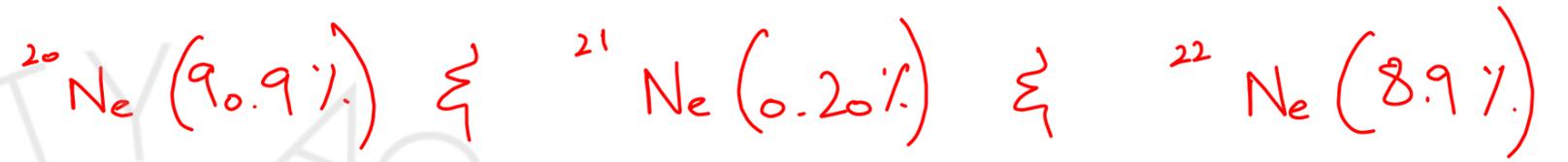
$$108 = 18x$$

$$x = 6$$

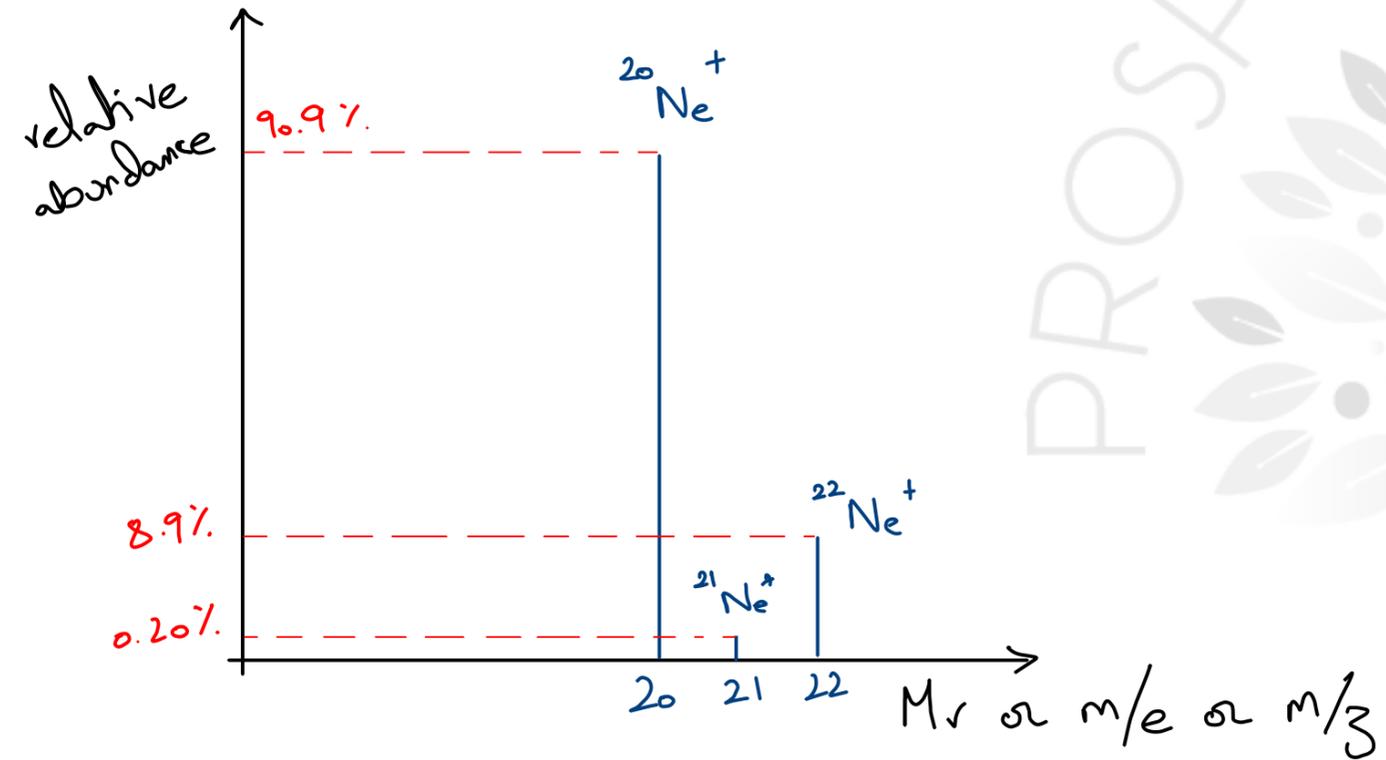
6 [1]

Mass Spectrometry:-

Neon has 3 naturally occurring isotopes:-



The mass spectrum of Neon looks like this:-

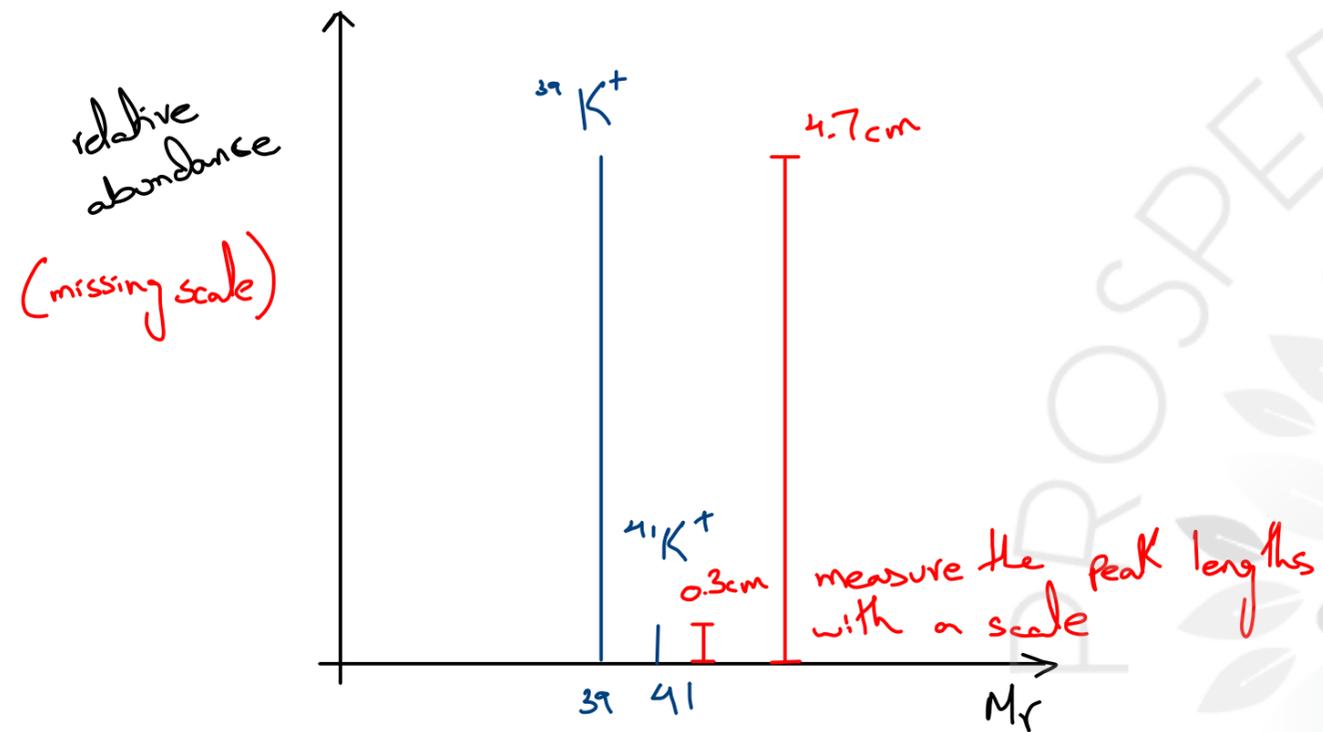


The Ar of Neon can be found out:-

$$\begin{aligned} A_r &= \left(\frac{90.9}{100} \times 20 \right) + \left(\frac{0.20}{100} \times 21 \right) + \left(\frac{8.9}{100} \times 22 \right) \\ &= 20.2 \end{aligned}$$

- We only give +1 charge so it can be assumed that any value on the x axis is the mass number.

Q. The mass spectrum of potassium is shown:-



a) What is the relative abundance of the ^{39}K and ^{41}K ?

$$\% \text{ of } ^{39}\text{K} = \frac{4.7}{3 + 4.7} \times 100 = 94\%$$

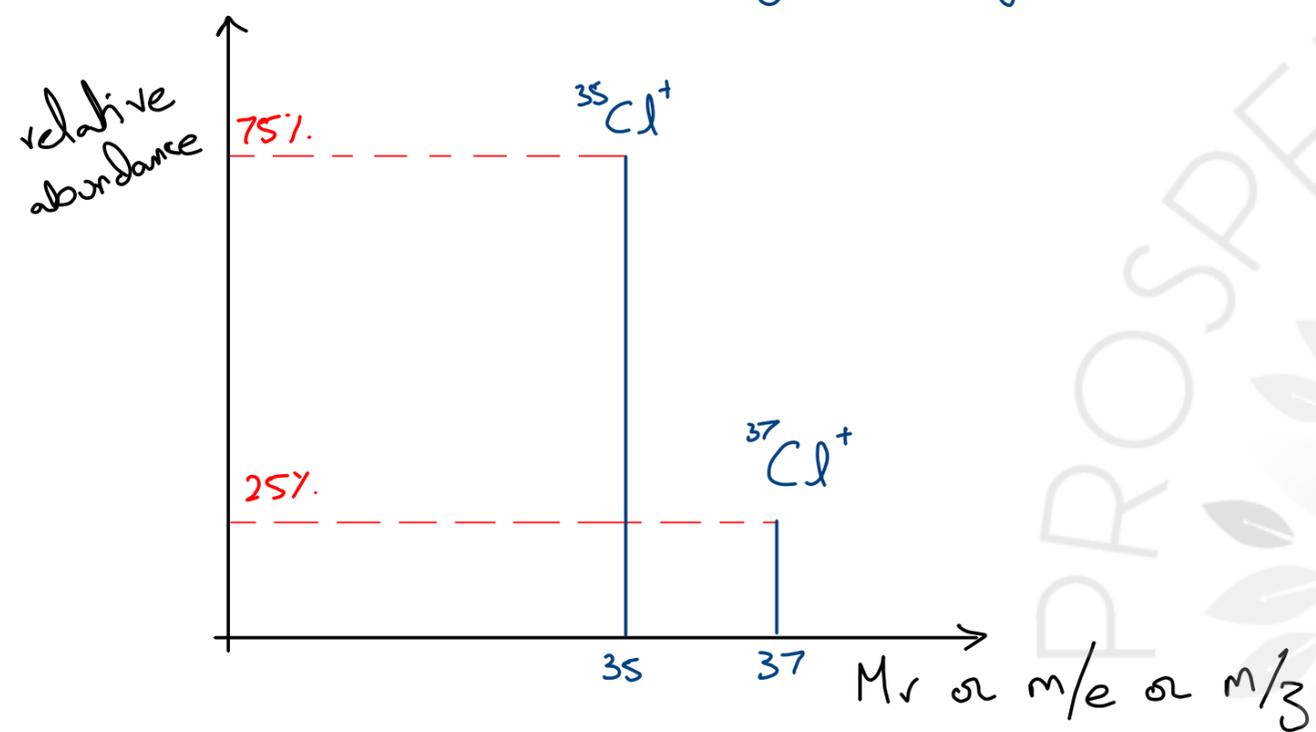
$$\% \text{ of } ^{41}\text{K} = \frac{0.3}{4.7 + 0.3} = 6\%$$

b) What is the A_r of potassium?

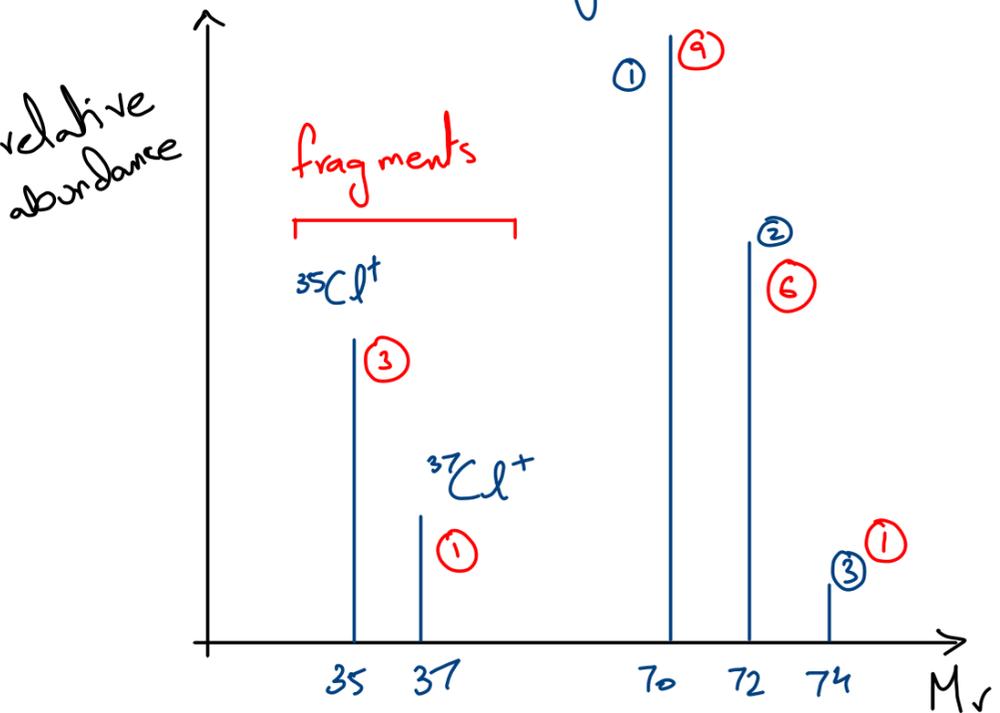
$$A_r = \left(\frac{94}{100} \times 39 \right) + \left(\frac{6}{100} \times 41 \right) = 39.12 \approx 39.1$$

Mass spectrum of Cl and Cl₂ :- (Learn)

- Chlorine has 2 naturally occurring isotopes :- ³⁵Cl (75%) and ³⁷Cl (25%)



- The mass spectrum of Cl₂ looks like this:-



$$1) \left[{}^{35}\text{Cl} - {}^{35}\text{Cl} \right]^+ = \frac{3}{4} \times \frac{3}{4} = \frac{9}{16}$$

$$2) \left[{}^{35}\text{Cl} - {}^{37}\text{Cl} \right]^+ \text{ or } \left[{}^{37}\text{Cl} - {}^{35}\text{Cl} \right]^+ = \left(\frac{3}{4} \times \frac{1}{4} \right) + \left(\frac{1}{4} \times \frac{3}{4} \right) = \frac{6}{16}$$

$$3) \left[{}^{37}\text{Cl} - {}^{37}\text{Cl} \right]^+ = \frac{1}{4} \times \frac{1}{4} = \frac{1}{16}$$

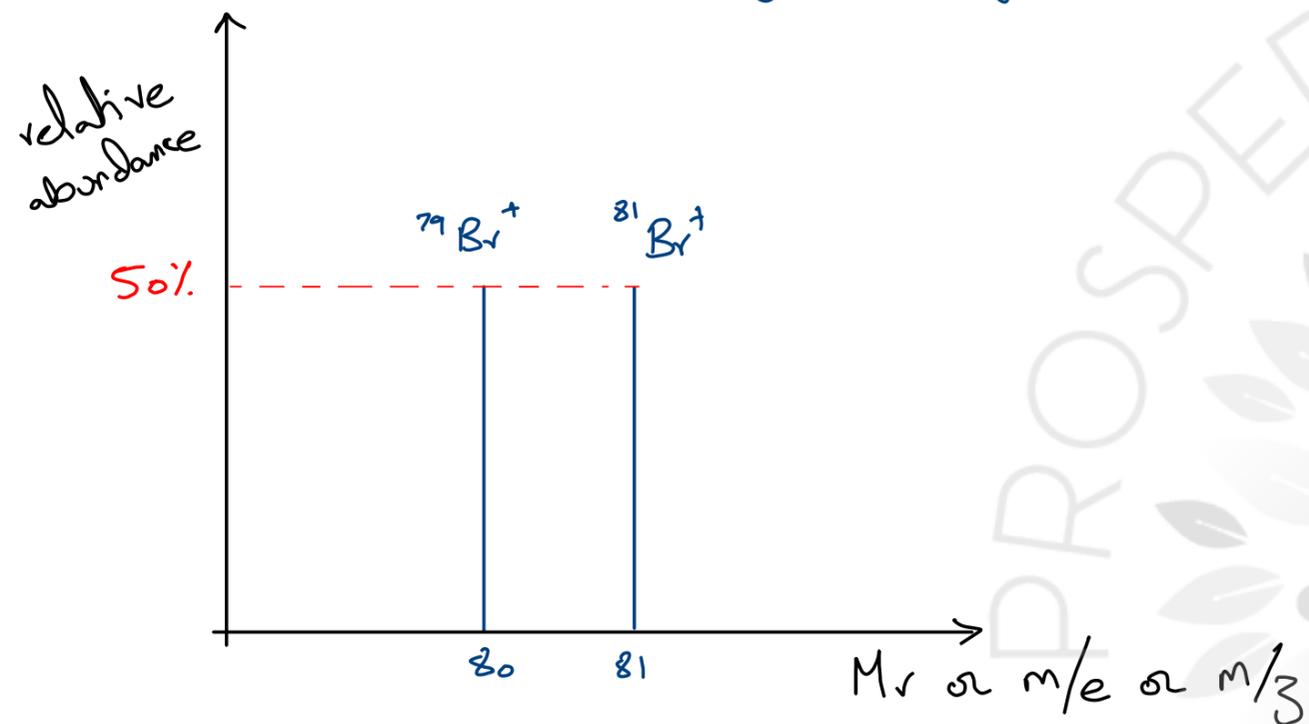
The molecules are in the ratio 9 : 6 : 1

The fragments are in the ratio 3 : 1

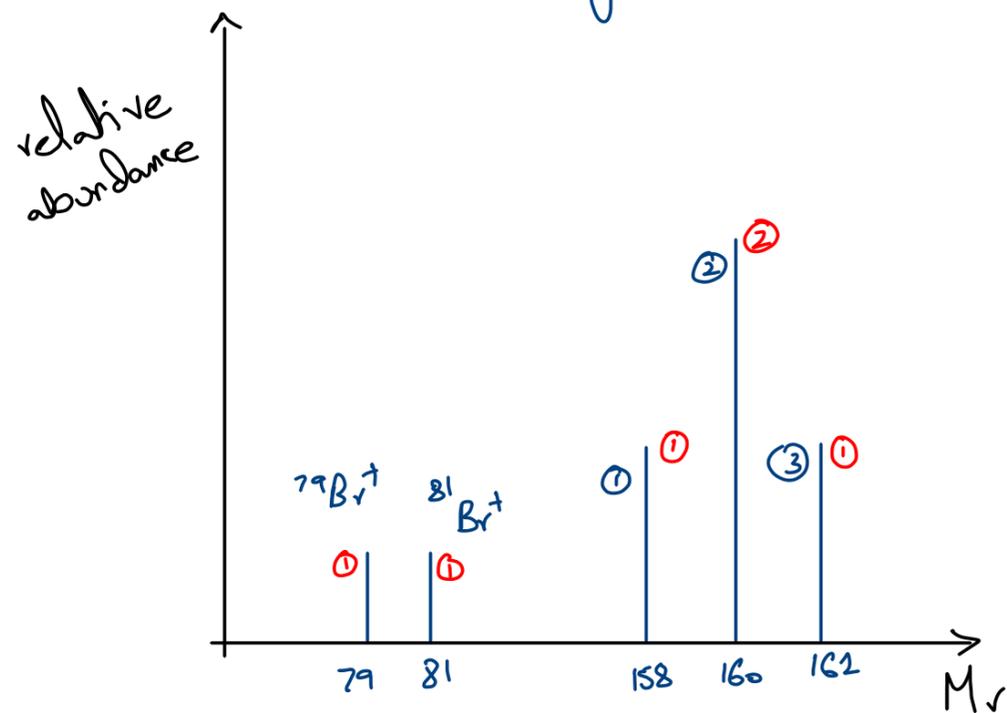
There is no ratio between fragments and molecules

Mass spectrum of Br and Br₂ :- (Learn)

- Bromine has 2 naturally occurring isotopes :- ⁷⁹Br (50%) and ⁸¹Br (50%)



- The mass spectrum of Br₂ looks like this:-



$$1) \left[{}^{79}\text{Br} - {}^{79}\text{Br} \right]^+ = \frac{1}{2} \times \frac{1}{2} = \frac{1}{4}$$

$$2) \left[{}^{79}\text{Br} - {}^{81}\text{Br} \right]^+ \text{ or } \left[{}^{81}\text{Br} - {}^{79}\text{Br} \right]^+ = \left(\frac{1}{2} \times \frac{1}{2} \right) + \left(\frac{1}{2} \times \frac{1}{2} \right) = \frac{1}{2}$$

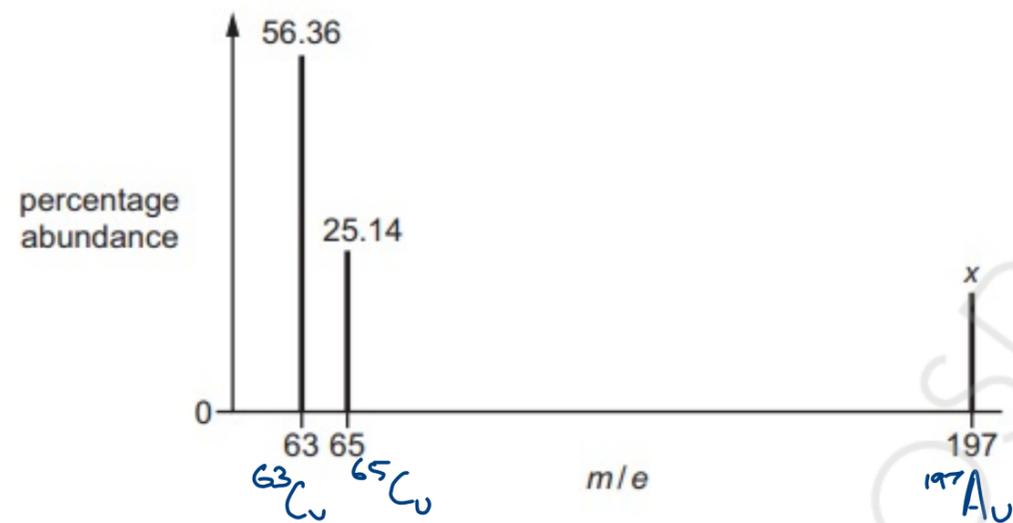
$$3) \left[{}^{81}\text{Br} - {}^{81}\text{Br} \right]^+ = \frac{1}{2} \times \frac{1}{2} = \frac{1}{4}$$

The molecules are in the ratio 1 : 2 : 1

The fragments are in the ratio 1 : 1

There is no ratio between fragments and molecules

(d) Tumbaga is an alloy of copper and gold. A sample of tumbaga was analysed. The mass spectrum of the sample is shown.



(i) Calculate the percentage abundance of gold, x , in the sample of tumbaga.

$$100 = 56.36 + 25.14 + x$$

$$x = 18.5$$

$$x = 18.50 \% \quad [1]$$

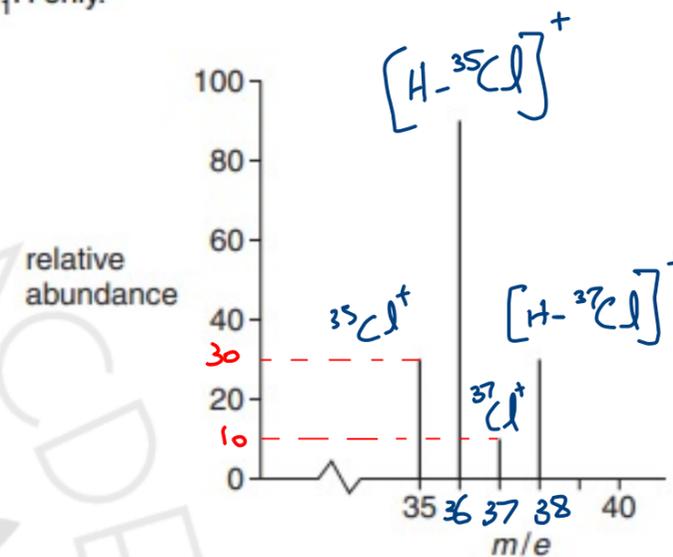
(ii) Calculate the relative atomic mass, A_r , of the copper present in this sample. Give your answer to **two** decimal places.

$$A_r = \left(\frac{56.36}{56.36 + 25.14} \times 63 \right) + \left(\frac{25.14}{56.36 + 25.14} \times 65 \right)$$

$$= 63.62$$

$$A_r(\text{Cu}) = 63.62 \quad [2]$$

(b) In a mass spectrometer some hydrogen chloride molecules will split into atoms. The mass spectrum of HCl is given. Chlorine has two isotopes. The hydrogen involved here is the isotope ^1H only.



(i) What particle is responsible for the peak at mass 35? $^{35}\text{Cl}^+$

(ii) What particle is responsible for the peak at mass 38? $[\text{H}-^{37}\text{Cl}]^+$ [2]

(c) Use the relative heights of the peaks to determine the proportions of the two isotopes of chlorine. Explain simply how you obtained your answer.

$$\frac{30}{30+10} \times 100 = 75\%$$

$^{35}\text{Cl}^+$

$$\frac{10}{30+10} \times 100 = 25\%$$

$^{37}\text{Cl}^+$

[2]

(d) Use your answer to (c) to explain why chlorine has a relative atomic mass of 35.5.

$$A_r = \left(\frac{75}{100} \times 35 \right) + \left(\frac{25}{100} \times 37 \right) = 35.5$$

[1]

- 2 Oxygen has three stable isotopes, ^{16}O , ^{17}O and ^{18}O . All three isotopes are present in a sample of oxygen gas, O_2 , which was analysed using a mass spectrometer.

How many peaks associated with the O_2^+ ion would be expected?

A 3

B 5

C 6

D 9

